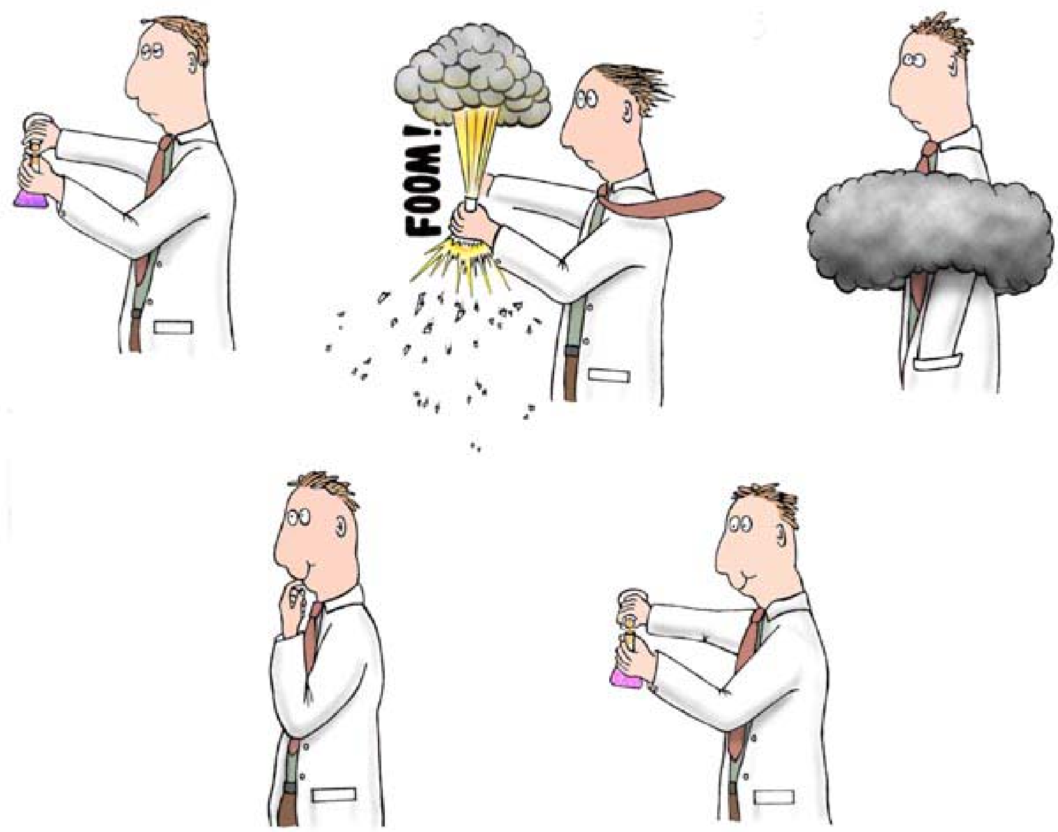
Chemistry I

Chemical Bonding
How It All Goes Together



Chemical Bonding

Bonding in any element will take place with only the ________________________________
The valence shell electrons are found in the____

By looking at the _________________________and/or the _________________________one is able to identify these valence electrons.

Chemical bonding occurs:

So atoms can achieve maximum stability:

1. ____
2. ____
3. ____

By opposite charges______________________, Like charges ______________________

The Octet Rule

The **octet rule** says that **atoms** tend to **gain, lose or share electrons** so as to have

____

BUT there are **many bonding situations where it does** ______________________________

Exceptions to the Octet Rule

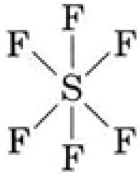
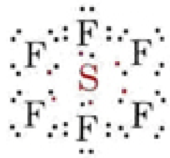
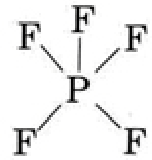
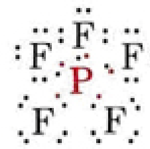
1. Elements in the third period and below can accommodate more than an octet of electrons. Elements **such as** ____**under some circumstances form more bonds than the rule allows.**
2. Some stable molecules simply____.

   This usually occurs in compounds containing ____or_____.
3. Free Radicals

Exceptions to the Octet Rule: Expansion

Atoms, which have room for more than 8 electrons in their outer shell, may form bonds, which result in____
(______________numbers are stable)

Examples: SF_6 and PF_5

 or   or 

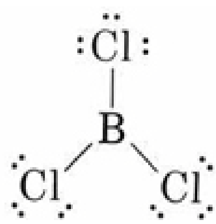
Exceptions to the Octet Rule: when there are not enough

Examples of this exception are provided by $BeCl_2$ and BCl_3.

Since Cl atoms do not readily form multiple bonds, we expect the Be atom to be joined to each Cl atom by a single bond.

Instead of an octet the valence shell of Be contains only ___________electron pairs.

Similar arguments can be applied to boron trichloride, BCl_3, which looks like this:



The valence shell of boron has only ______________pairs of electrons.

Molecules such as $BeCl_2$ and BCl_3 are referred to as **electron deficient** because some atoms do not have ____

Exceptions to the Octet Rule: Free Radicals

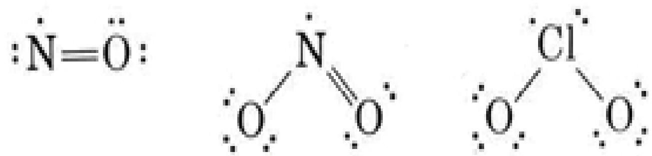
There are a few stable molecules which contain an odd number of electrons.

These molecules, called ***"free radicals"***, contain at least one unpaired electron, a clear violation of the octet rule.

Free radicals play many important roles a wide range of chemistry fields.

Three well-known examples of such molecules are nitrogen (II) oxide, nitrogen (IV) oxide, and chlorine dioxide.

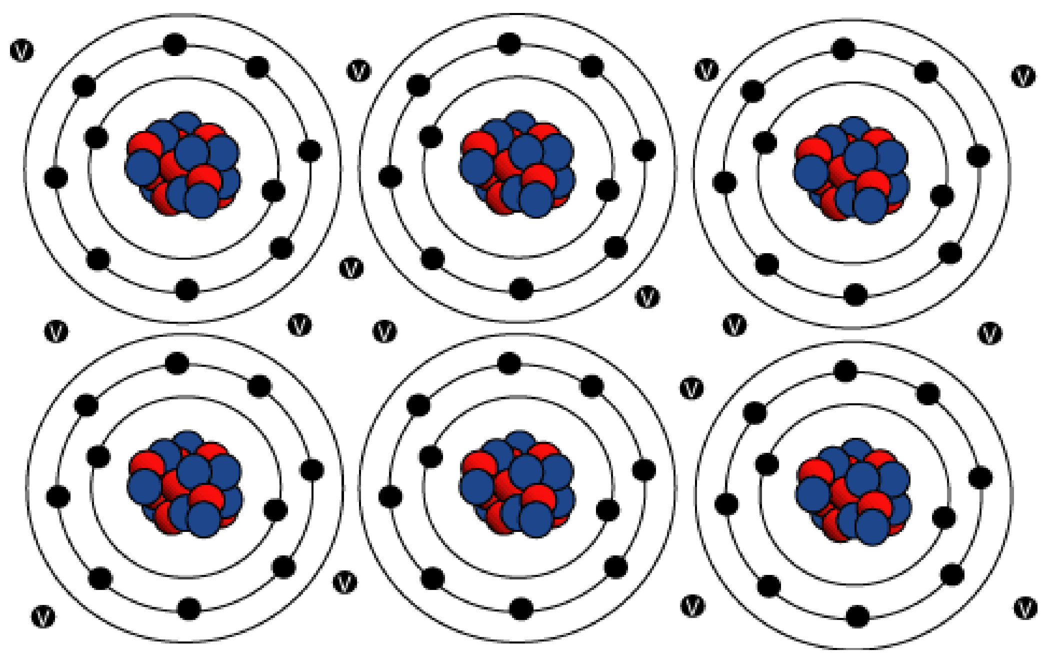
The most plausible Lewis structures for these molecules are:



Metallic bonding

1. Occurs between atoms of the ______________________________
2. Atoms have 1,2 or 3 valence electrons, therefore there are many *vacancies in valence shell.*
3. Nuclei of the atoms are in a ______________________________
4. When electron clouds overlap ______________________________

____



Properties of substances with metallic bonding

Property	Reason
Metals are dense	
Metals have high melting and boiling points	
Metals are good conductors of heat	
Metals are good conductors of electricity	
Metals are malleable and ductile	
Metals are lustrous (shiny)	

Ionic Bonding

Ionic bonding occurs between *metals* and *non-metals*.

A. ____________________atoms have a *low number of valence electrons* and a *low electronegativity.*

B. ___________________________________atoms have *numerous valence electrons*.

If the electron clouds overlap (bond)..

Metals

1. lose valence electrons
2. achieve a stable valence shell (usually 8 e-)
3. will gain a ______________charge, and becomes_______________________

Non-metals

1. gain valence electrons
2. achieve a stable valence shell (usually 8 e-)
3. Will gain a _______________charge and becomes________________________

Formation of positive ions (cations)

The charge is the number of valence e^{-1} 's it has to_______________.

1) Na atom [Ne] $3s^1$ loses one electron to become a sodium ion (Na^{+1}) with [He] $2s^22p^6$configuration

2) K atom [Ar] $4s^1$ loses one electron to become a potassium ion (K^{+1}) with [Ne] $3s^23p^6$ configuration.

This is common for ALL ___________________________

3) Mg atom [Ne]$3s^2$ loses 2 electrons to become a magnesium ion (Mg^{+2}) with [He] $2s^22p^6$ configuration

4)Ca atom [Ar] $4s^2$ loses 2 electrons to become a calcium ion (Ca^{+2}) with [Ne] $3s^23p^6$ configuration.

This is common for ALL ___________________________

5) Al atom [Ne] $3s^23p^1$ **loses 3 electrons** to become an aluminum ion (Al^{+3}) with [He] $2s^22p^6$ configuration.

This is common for **Group 13 metals ONLY**

General note: metals will always make __________________________________

Formation of negative ions (anions)

1) F atom [He] $2s^22p^5$ goes to the fluoride ion (F^{-1}) with [He] $2s^22p^6$
2) Cl atom [Ne]$3s^23p^5$ goes to the chloride ion (Cl^{-1}) with [Ne] $3s^23p^6$
3) O atom [He] $2s^22p^4$ goes to the oxide ion (O^{-2}) with [He] $2s^22p^6$
4) The charge is the number of electrons needed to get 8 in the outermost level and make the charge negative.

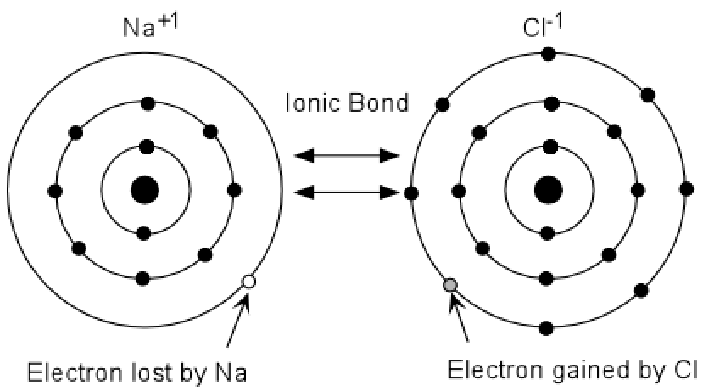
   Ex: O gains 2 e^- and has –2 charge

   1. Nonmetals will **usually** make ____________________________

Formation of ionic compounds

The positive and negative ions will attract each other to form a three dimensional continuous _______________________________(or crystal).

- Each positive ion is surrounded by a number of _______________________
- Each negative ion is surrounded by a number of_______________________
- The ratio of positive to negative ions in the lattice is determined by the charges of the ions



Properties of Ionic compounds and why they have these properties

Property	Reason
Melting point and boiling point are high	
Electrical conductivity is poor	
The crystals of ionic compounds are hard	
The crystals of ionic compounds are hard	

Covalent Bonding

is between ____

- Therefore all atoms included have *fairly high electronegativity* and *few vacancies in valence energy levels*.
- When they bond, they gain electrons to achieve ___________________ configuration.
- Hence, **electrons are** ___________________________________.

Sharing produces low energy (stable) electron arrangements that are isoelectronic (same # e^{-1} as) with the_________________________________

ie:

- Full outer shell (eg He $1s^2$; Ne $1s^2\ 2s^2\ 2p^6$)

  8 electrons (4 pairs) in outer shell (eg Ar $1s^2\ 2s^2\ 2p^6\ 3s^2 3p^6$)

When more than one pair of electrons are shared

1. There some situations in covalent bonding where there are insufficient electrons.
2. When this occurs, atoms will share______________________________

Carbon is one of the elements that does this quite readily.

If two atoms **share1pair** of electrons, the bond is called a________________**bond.**

If two atoms **share 2 pairs** of electrons, the bond is called a ______________**bond.**

If two atoms **share 3 pairs** of electrons, the bond is called a ______________**bond.**

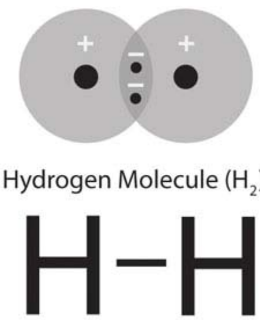
Covalence is the number of electrons an atom needs to produce a stable outer shell.(how many does it need to get to meet the octet rule.)

The number of *shared pairs (covalent bonds)* of electrons an atom forms.

eg Hydrogen *H* needs 1 additional electron therefore the covalence is 1

Covalent bonding in the hydrogen molecule

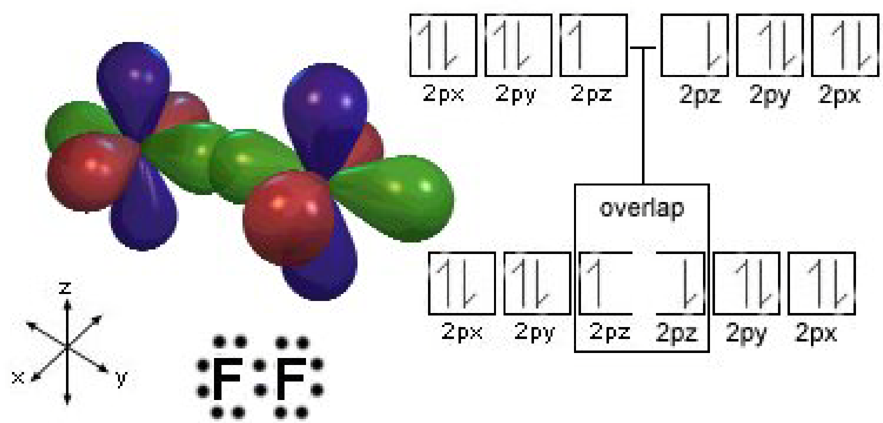
When hydrogen bonds it bonds with 1 pair of e-'s shared between 2 atoms – a *covalent bond*.



Hydrogen molecule consists of 2 covalently bonded hydrogen atoms, which have no tendency to bond further (both have achieved a stable outer shell).
Each molecule exists independently

Covalent bonding in other diatomic molecules

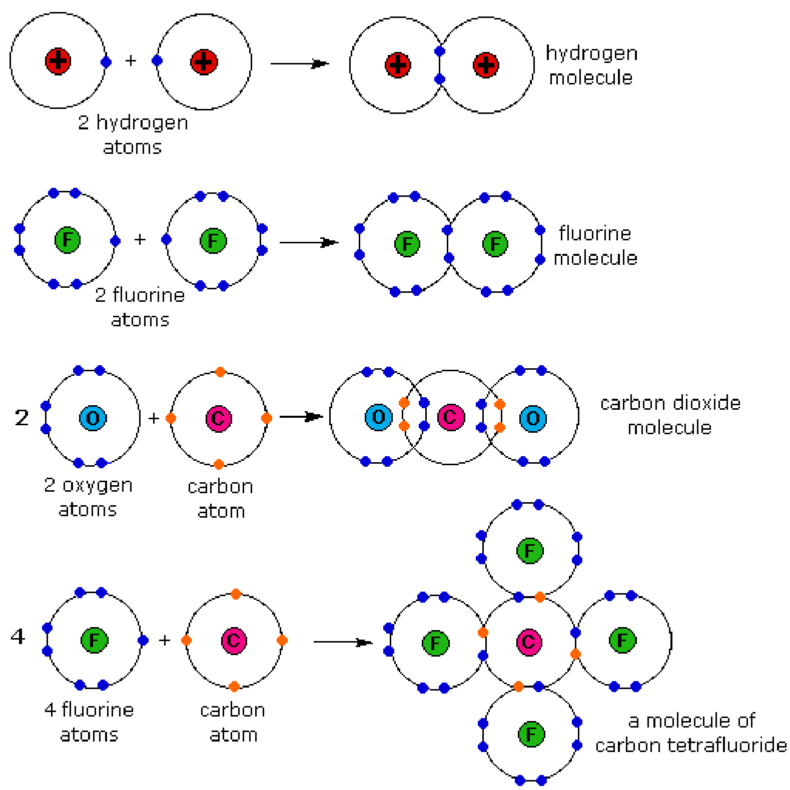
Fluorine F [He] $2s^22p^5$ needs one electron to achieve Noble gas configuration.



Properties of covalently bonded compounds

Property	Reason
Do not conduct electricity	
Melting and boiling points are low	
Molecules are soft	
Covalent compounds aren't usually very soluble in water	

Covalent Bonding Examples



Not ALL electrons are shared equally in a covalent bond

In covalent bonds, where the electronegativities are equal between the two atoms, the atoms share electron pair equally are called ____

In covalent bonds, where the electronegativities are not equal, the atom with the higher electronegativity will have a greater pull on the shared electron pair.

These are called ___________________________________

One of the atoms in this bond "hogs" the shared electron pair.

(usually the most electronegative atom).

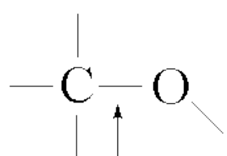
The atom tending to "hog" the electron pair acquires a slight __________________*charge*

The atom tending to "lose" electron pair acquires a slight ______________________*charge*

The difference between C and O electronegativities is 1.0 Pauling unit.

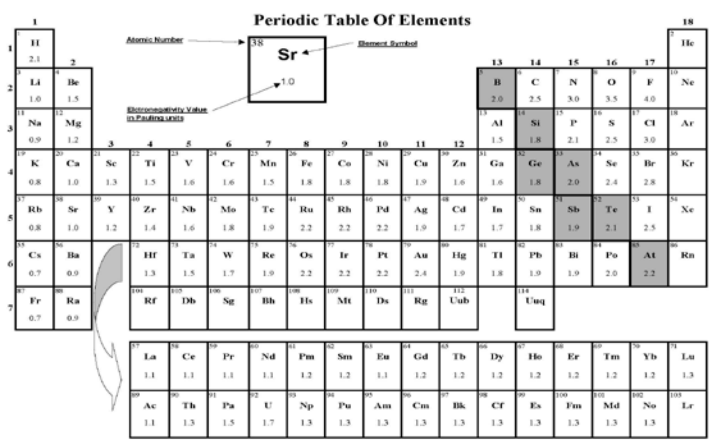
(3.5 – 2.5 = 1.0)

Therefore the bond between these 2 atoms is polar covalent.



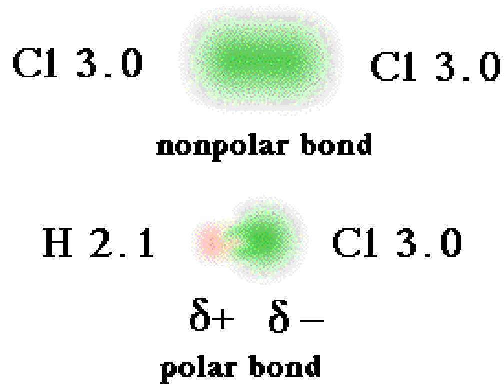
Polar Covalent Bond

The bond *is* called a **polar covalent bond.**



bond type	electronegativity difference
nonpolar covalent	0.0-0.3
polar covalent	0.4-1.6
ionic	1.7-3.2

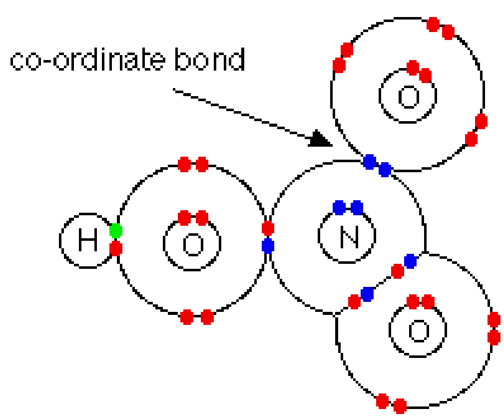
Covalent Bonds

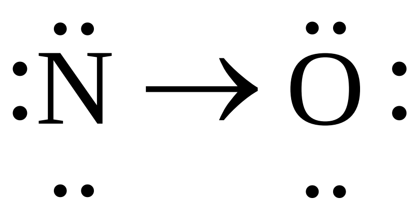


Co-ordinate Bonds

A rare covalent bond that occurs when the **shared pair** of e-'s in a covalent bond is **provided by 1 atom**.

It is represented by an arrow, showing the direction in which the shared e-'s are provided. Co-ordinate bonds aren't very common. They usually occur in atoms that cannot expand the octet.





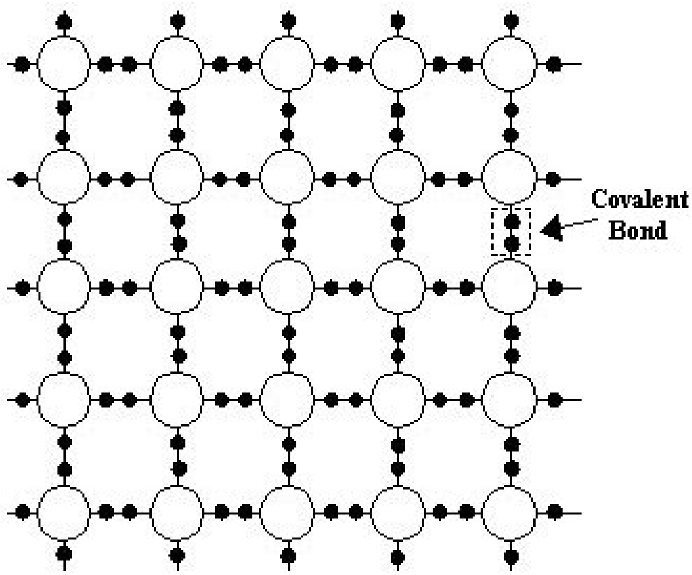
Covalent Network (Lattice)

Bonding between_______________________________.

Form covalent bonds between the atoms (shared electron pairs).

Do not form **separate (discrete) molecules** but a ***continuous network***.

Examples: diamonds and quartz crystals



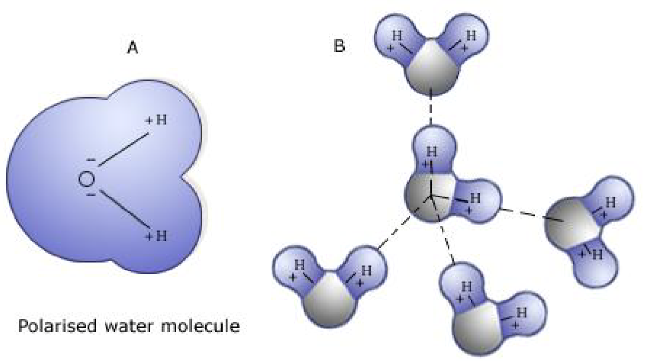
Properties of covalent network substances.

Property	Reason
Poor Electrical conductivity	
Very high Melting points	
They are hardness	.
Covalent network substances are brittle.	

Intermolecular Forces Covalent Compounds

Force	Molecule type	Explanation
Covalent Compounds		
London		
Dipole-dipole		
Hydrogen bonding		
Ionic Compounds & Metals		
electrostatic		

Hydrogen Bonding



(A) Polarised covalent bonds link the hydrogen and oxygen atoms in a water molecule. (B) Hydrogen bonds between adjacent water molecules. Hydrogen bonds are represented in diagrams by dashed or dotted lines. and covalent bonds by solid lines.

Molecular Geometry (shape)

The way that some atoms covalently bond together will cause them to have a particular molecular shape.

There are three theories that attempt to explain this:

- VESPER theory
- Valence Bond Theory
- Molecular orbital theory

VESPER Theory (the one we use)

The valence shell electron-pair repulsion model (VESPR) was devised to account for these__________________________________.

In this model, atoms and pairs of electrons will be ________________________________

____

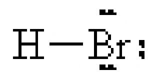
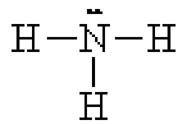
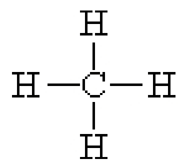
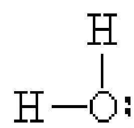
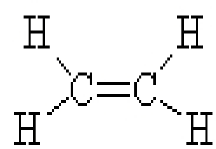
A simple model for demonstrating the behaviour of electron pairs under the influence of their mutual repulsion is provided by a set of spherical balloons of equal size.

Four balloons tied together so that they squeeze each other fairly tightly, they inevitably adopt the *tetrahedral* arrangement shown for CH_4 in Table 1.

Although it is possible to flatten the balloons on a table until they are all in the same plane, they invariably spring back to the tetrahedral configuration as soon as the pressure is removed.

A similar behaviour is found if two, three, five, or six balloons are tightly tied together, except that in each case a different stable shape is adopted once the balloons are left to themselves.

Table 1

Formula	Lewis Structure	Molecular Geometry
HBr		linear
NH_3		pyramidal
CH_4		tetrahedral
H_2O		bent
C_2H_4		triangular

The solution to the Schrodinger Equation provides for the following atomic orbitals:

1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 3*d*, 4*s*, 4*p*, 4*d*, 4*f*, etc.

An atomic orbital is really the energy state of an electron bound to an atomic nucleus. The energy state changes when one atom is bonded to another atom. The way this works is by combining the orbitals to give new orbitals of equal energy. This is called the **hybridization** of atomic orbitals.

Bottom line: we can say that an imaginary mixing process converts a set of atomic orbitals to a new set of ____or **hybrid orbitals**.

At this level, we consider the following hybrid orbitals:

a) *sp*
b) sp^2
c) sp^3

HYBRIDIZATION OF CARBON

The element, carbon, is one of the ______________________________ ______________terms of the number of compounds it may form. It may form virtually an infinite number of compounds.

Since both the_______and the ______ sublevels are half-filled, the excited state is relatively stable.

This is largely due to the types of bonds it can form and the number of different elements it can join in bonding.

Carbon may form ________________________&____________bonds. The hybridization of carbon involved in each of these bonds will be investigated in our laboratory exercise.

From the ground state electron configuration, one can see that carbon has four valence electrons, two in the 2s subshell and two in the 2p subshell.

$1s^2 2s^2 2p^2$ or $\underset{1s}{\underline{\uparrow\downarrow}}\ \underset{2s}{\underline{\uparrow\downarrow}}\ \underset{2p_x}{\underline{\uparrow}}\ \underset{2p_y}{\underline{\uparrow}}\ \underset{2p_z}{\underline{\quad}}$

The 1s electrons are considered to be ________________________and are not available for bonding.

Carbon will form an excited state by promoting one of its 2s electrons into its empty 2p orbital and hybridize from the excited state.

$\underset{2s}{\underline{\uparrow}}\ \underset{2p_x}{\underline{\uparrow}}\ \underset{2p_y}{\underline{\uparrow}}\ \underset{2p_z}{\underline{\uparrow}}$

By forming this excited state, carbon will be able to form four bonds. The excited state configuration is said to be ***sp^3 hybridized.***

Let's look at an example of each of the hybridizations of carbon.

For our first example, let's choose methane, CH_4 . Draw the Lewis structure.

This Lewis structure shows ________**groups around the**________________ **atom**. This means four hybrid orbitals have formed. In order to form four hybrid orbitals, four atomic orbitals have been mixed. The s orbital and all three p orbitals have been mixed, thus **the hybridization is _________.**

For our first example, let's choose ethylene, $C_2 H_4$ Draw the Lewis structure.

This Lewis structure shows ________**groups around the carbon atom**. This means three hybrid orbitals have formed. In order to form four hybrid orbitals, four atomic orbitals have been mixed. The s orbital and all two p orbitals have been mixed, thus **the hybridization is** ______________

For our first example, let's choose acetylene, C_2 H_2 Draw the Lewis structure.

This Lewis structure shows__________ **groups around the carbon atom.** This means two hybrid orbitals have formed. In order to form this bond, two atomic orbitals have been mixed. The s orbital and one p orbital have been mixed, thus **the hybridization is ______________________**

Summary

The number of groups represents how many hybrid orbitals have formed. The number of hybrid orbitals formed equals the number of atomic orbitals mixed. The description of the atomic orbitals mixed is equivalent to the hybridization of the carbon atom.

Hybridization	Orbitals involved	Sites for bonding
sp^3		
sp^2		
sp		

Resonance

There are a number of compounds and polyatomic ions that cannot be written using

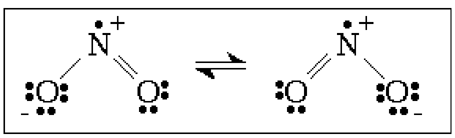
____

Linus Pauling developed what today is called "resonance theory."

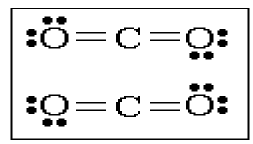
Resonance happens when more than one valid Lewis dot-diagram can be written for a molecule or ion.

When this happens, the true structure is a blend of all the different possible structures.

Here is another example, using the molecule CO_2:



Carbon Dioxide Resonance Structures



OK, here's the deal. Neither one of those two structures really does exist.

The real molecule that exists in nature is a "resonance hybrid" between the two.

The real molecule acts as if it had one and one-half bonds between each of the two structures. Here is the Lewis structure possibilities for N_2O (see above).